Chem 2 AP Homework #4-3: Problems pg. 155 # 4.76 – 4.80

76 **Distilled water must be used in the gravimetric analysis of chlorides. Why?**
Because tap water contains chloride ions that can interfere with the tests.

77 **If 30.0 mL of 0.150 M $CaCl_2$ is added to 15.0 ml of 0.100 M $AgNO_3$, what is the mass in grams of AgCl precipitate?**

$$\underset{\text{30.0 mL of 0.150 M}}{CaCl_2(aq)} + \underset{\text{15.0 mL of 0.100 M}}{2AgNO_3(aq)} \longrightarrow Ca(NO_3)_2(aq) + \underset{\text{? g}}{2AgCl(s)}$$

We need to determine the limiting reagent. Ag^+ and Cl^- combine in a 1:1 mole ratio to produce AgCl. Let's calculate the amount of Ag^+ and Cl^- in solution.

$$\text{mol } Ag^+ = 15.0 \text{ mL soln} \times \frac{0.100 \text{ mol } Ag^+}{1000 \text{ mL soln}} \times = 1.50 \times 10^{-3} \text{ mol } Ag^+$$

$$\text{mol } Cl^- = 30.0 \text{ mL soln} \times \frac{0.150 \text{ mol } CaCl_2}{1000 \text{ mL soln}} \times \frac{2 \text{ mol } Cl^-}{1 \text{ mol } CaCl_2} = 9.00 \times 10^{-3} \text{ mol } Cl^-$$

Since Ag^+ and Cl^- combine in a 1:1 mole ratio, $AgNO_3$ is the limiting reagent. Only 1.50×10^{-3} mole of AgCl can form. Converting to grams of AgCl:

$$1.50 \times 10^{-3} \text{ mol AgCl} \times \frac{143.4 \text{ g AgCl}}{1 \text{ mol AgCl}} = \boxed{\textbf{0.215 g AgCl}}$$

78 **A sample of 0.6760 g of an unknown Ba^{2+} compound is treated with excess Na_2SO_4 to form 0.4105 g of $BaSO_4$ precipitate. What is the mass % of Ba in the original compound?**
Let's first map out what we know, what we want, and what we need to get there:

want to calculate → mass % Ba; need to find → mass Ba; given → mass of sample

$$\text{mass \% Ba} = \frac{\text{mass Ba}}{\text{mass of sample}} \times 100\%$$

First, we calculate the mass of Ba in 0.4105 g of the $BaSO_4$ precipitate:

$$\text{? mass of Ba} = 0.4105 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.4 \text{ g } BaSO_4} \times \frac{1 \text{ mol Ba}}{1 \text{ mol } BaSO_4} \times \frac{137.3 \text{ g Ba}}{1 \text{ mol Ba}}$$

$$= 0.2415 \text{ g Ba}$$

Next, we calculate the mass percent of Ba in the unknown compound.

$$\textbf{\%Ba by mass} = \frac{0.2415 \text{ g}}{0.6760 \text{ g}} \times 100\% = \boxed{\textbf{35.72\%}}$$

79 How many grams of NaCl are required to precipitate most of the Ag^+ ions from 2.50×10^2 mL of 0.0113 M $AgNO_3$ solution? Write the net ionic equation for the reaction.

The net ionic equation is: $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

One mole of Cl^- is required per mole of Ag^+. First, find the number of moles of Ag^+.

$$\text{mol } Ag^+ = 2.50\times10^2 \text{ mL soln} \times \frac{0.0113 \text{ mol } Ag^+}{1000 \text{ mL soln}} = 2.83\times10^{-3} \text{ mol } Ag^+$$

Now, calculate the mass of NaCl using the mole ratio from the balanced equation.

$$2.83\times10^{-3} \text{ mol } Ag^+ \times \frac{1 \text{ mol } Cl^-}{1 \text{ mol } Ag^+} \times \frac{1 \text{ mol NaCl}}{1 \text{ mol } Cl^-} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = \boxed{\textbf{0.165 g NaCl}}$$

80 The concentration of Cu^{2+} ions in the water discharged from a certain industrial plant is determined by adding excess Na_2S solution to 0.800 L of the water. The molecular equation is:

$$\textbf{Na}_2\textbf{S (aq) + CuSO}_4\textbf{ (aq)} \rightarrow \textbf{Na}_2\textbf{SO}_4\textbf{ (aq) + CuS (s)}$$

Write the net ionic equation and calculate the molar concentration of Cu^{2+} in the water sample if 0.0177 g of solid CuS is formed.

The net ionic equation is: $Cu^{2+}(aq) + S^{2-}(aq) \longrightarrow CuS(s)$

The answer sought is the molar concentration of Cu^{2+}, that is, moles of Cu^{2+} ions per liter of solution. The factor-label method is used to convert, in order:

$$\text{g of CuS} \rightarrow \text{moles CuS} \rightarrow \text{moles } Cu^{2+} \rightarrow \text{moles } Cu^{2+} \text{ per liter soln}$$

$$\textbf{mol } \textbf{Cu}^{2+} = 0.0177 \text{ g CuS} \times \frac{1 \text{ mol CuS}}{95.62 \text{ g CuS}} \times \frac{1 \text{ mol } Cu^{2+}}{1 \text{ mol CuS}} = \textbf{1.85}\times\textbf{10}^{-4} \textbf{ mol } \textbf{Cu}^{2+}$$

$$\textbf{[Cu}^{2+}\textbf{]} = \frac{1.85\times10^{-4} \text{ mol } Cu^{2+}}{0.800 \text{ L}} = \boxed{\textbf{2.31}\times\textbf{10}^{-4} \textbf{ M}}$$